EXTENSION ACTIVITY
GENERAL CHEMISTRY

RADII TRENDS KEY

Activity Directions

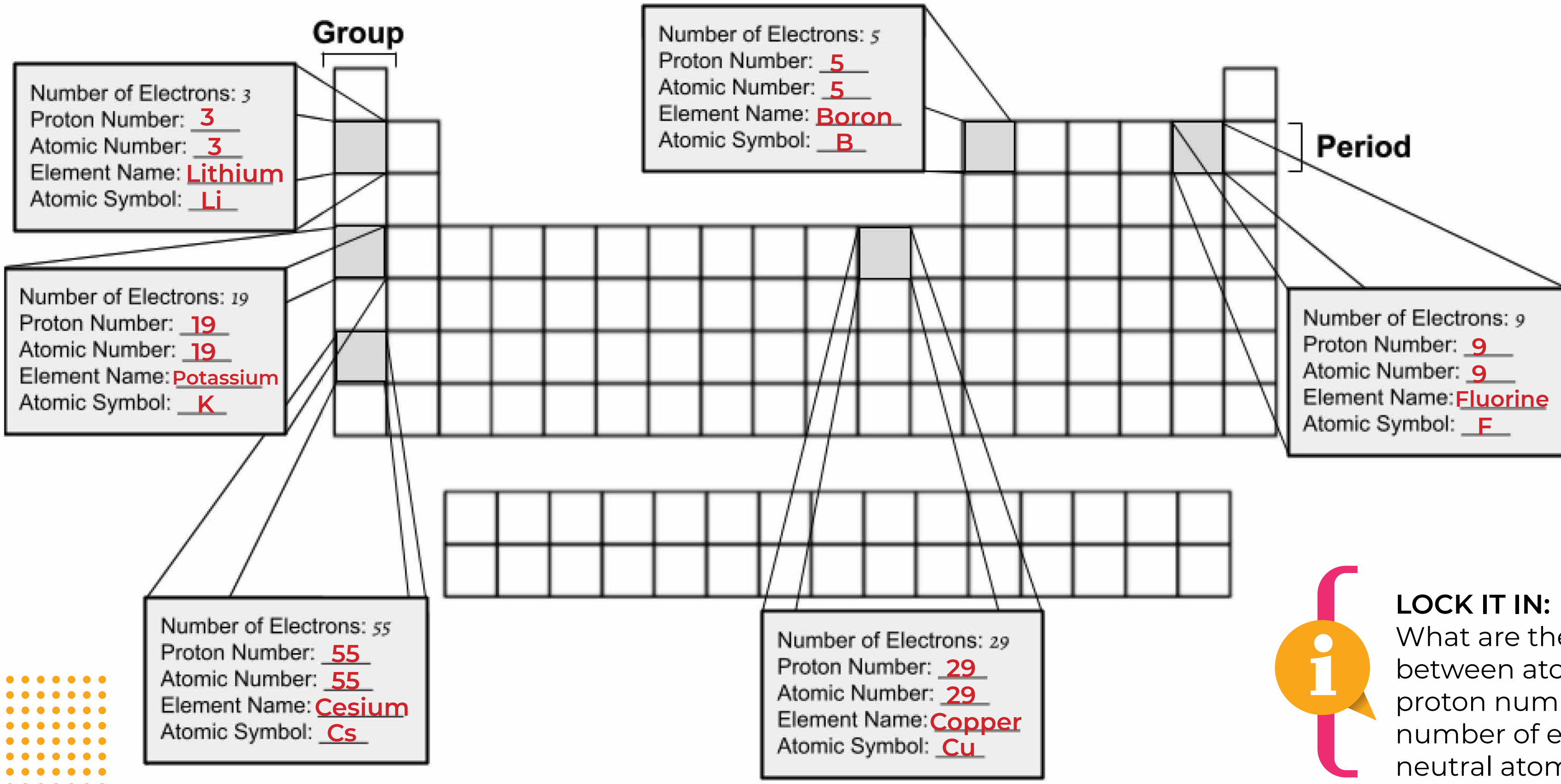
This activity will serve as practice for the topics covered in the Radii Trends game, as well as help you connect those topics to the periodic table and other associated concepts. This activity is best used in conjunction with not only the tutorial levels, but also supplementary learning resources such as course lectures, textbook reading, etc. Questions labeled "Lock It In" are simply opportunities for you to solidify what you have accomplished in each task and help ensure you meet each objective.

1. Log into Collisions and navigate to the Radii Trends Game.
2. Play the Tutorial levels, if you haven't done so already.
3. Exit the levels and enter the Radii Trends sandbox.
4. Follow all instructions as written below. Be sure to reference your course's textbook, lecture notes, etc. as needed.

OBJECTIVE 1

Demonstrate an understanding of the relationship between atomic number, proton number, the number of electrons, and the way these numbers increase on the periodic table.

TASK 1: For each of the elements indicated on the periodic table below, build the correct atom in the sandbox with the given number of electrons and then complete the missing data in the gray box.



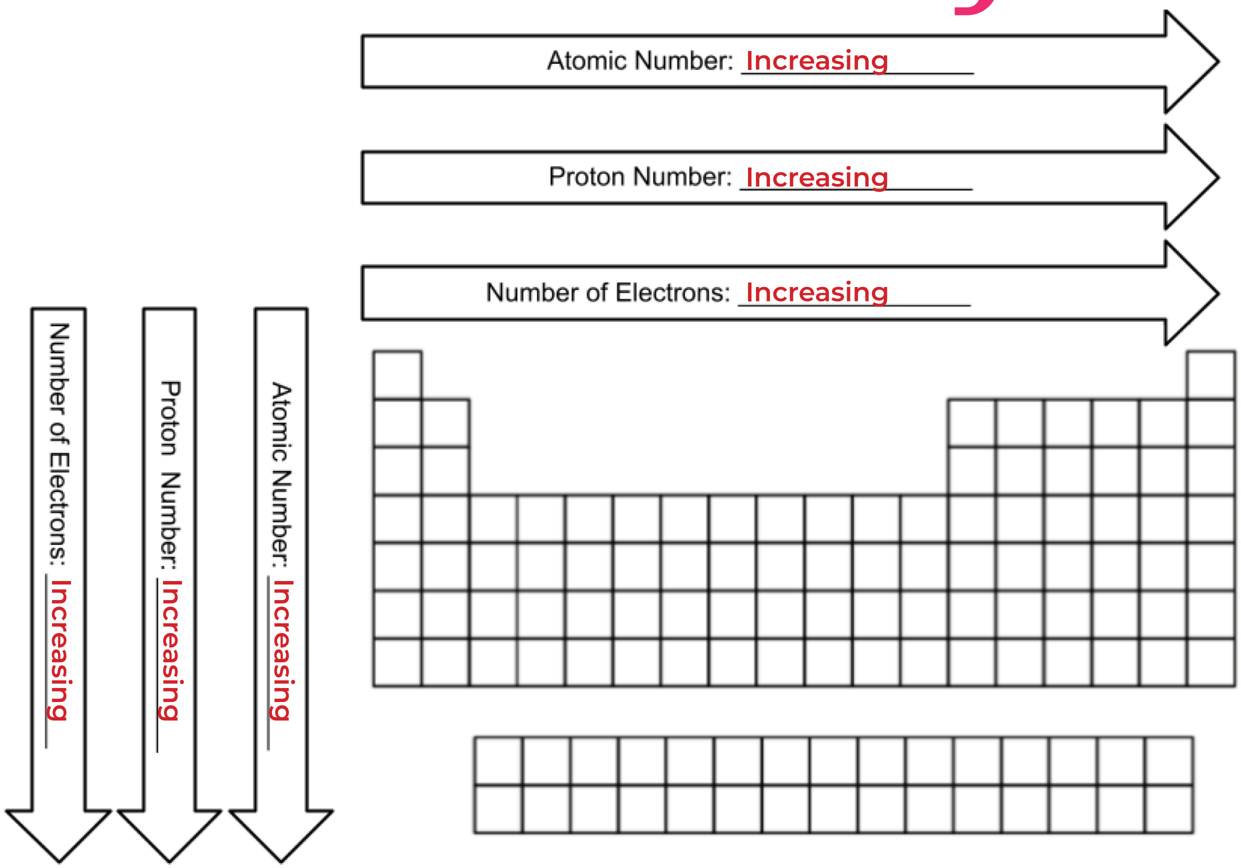
LOCK IT IN:
What are the relationships between atomic number, proton number, and the number of electrons in a neutral atom?

The atomic number of an element is the same as the number of protons in each of its atoms. In a neutral atom, the number of electrons equals the number of protons and thus is also the same as the atomic number.

OBJECTIVE 1

Demonstrate an understanding of the relationship between atomic number, proton number, the number of electrons, and the way these numbers increase on the periodic table.

LOCK IT IN:
Identify the relationship between atomic number, number of protons, and the number of electrons in a neutral atom as you go right across a period and down a group by filling in the arrows below with the terms “increasing” or “decreasing”.



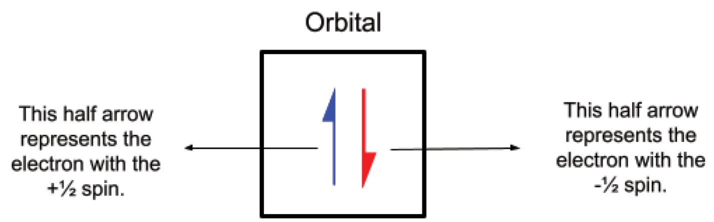
OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

A proper understanding of electrons and their organization requires knowledge of four key concepts— Heisenberg's uncertainty principle, the Pauli exclusion principle, aufbau principle, and Hund's Rule. A brief description of each is provided below as a reference before you begin the next task. However, it is important that you learn more about each in detail from your textbook or other resources.

- **HEISENBERG'S UNCERTAINTY PRINCIPLE:** concept that tells us that the more we know about the position of an electron, the less we can know about its velocity (and thus momentum). Without both such things, we do not know an electron's location. Keep this in mind as you place electrons in their orbitals in the sandbox.

- **PAULI EXCLUSION PRINCIPLE:** concept that tells us that no two electrons in an atom can have the same four quantum numbers. The first three quantum numbers easily distinguish most electrons from one another, however two electrons in the same orbital will share the first three quantum numbers in common. However, the Pauli Exclusion Principle is maintained because electrons in the same orbital will have opposite spin quantum numbers (+½ or -½). You will see this depicted in orbital diagrams by one electron in an orbital represented by a half arrow pointing up, while the other electron is represented by a half arrow pointing down. By convention the first electron to be placed into an orbital has the spin quantum number of +½.



- **AUFBAU PRINCIPLE:** concept that tells us that in an atom in its ground state, electrons fill the lowest available energy state before filling higher ones.

- **HUND'S RULE:** concept that tells us that due to the repulsive forces that exist within an orbital, electrons will fill them singly first before pairing up.

OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

TASK 2: Answer the questions below and complete the phosphorus activity.

Although the electrons you added to all of your atoms thus far seemed to be stationary, they are actually moving very quickly. Why can't we ever identify the exact location of an electron even though we can determine the location of other moving objects? Use **Heisenberg's uncertainty principle** as part of your answer.

Heisenberg's uncertainty principle tells us that the more we know about the position of an electron, the less we know about its velocity. The opposite is also true. This phenomenon is the result of the electron having the properties of both waves and particles simultaneously. Since identifying the location of an electron at any time would require fully knowing both its starting position and its velocity, we are never able to identify the location of an electron.

Create a phosphorus atom (Z = 15) and try skipping the first energy level by adding all of your electrons to the second energy instead. Explain why the **aufbau principle** prevented you from doing so.

The aufbau principle tells us that in an atom in its ground state, electrons fill the lowest available energy states before filling higher ones. When the phosphorus atom is built by putting electrons into the second energy level first, electrons are being given higher energy states before the lower ones have been filled. This action violates the aufbau principle.

Try creating the same phosphorus atom again. As you start to fill the 2p orbitals, put two electrons together in the first orbital. Explain how this violates **Hund's rule**.

Hund's rule tells us that due to the repulsive forces that exist within an orbital, electrons will fill orbitals singly first before pairing up. When the phosphorus atom is built by putting electrons together in orbitals before other orbitals of the same subshell have an electron, the action ignores that repulsive forces cause ground state atoms to have the highest number of partially filled orbitals with the same energy as possible before electrons pair up. Constructing the phosphorus atom in such a way violates Hund's rule.

OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

LOCK IT IN:
Write out the electron configuration and the noble gas configuration of **phosphorus**, and then fill out the orbital diagram using up and down arrows. Be sure to fill out the diagram keeping with the aufbau principle, Hund's rule, and the Pauli exclusion principle. **Be sure** to label each sublevel (one has already been labeled for you).

Phosphorus

Orbital Diagram: [↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑|↑|↑]

1s

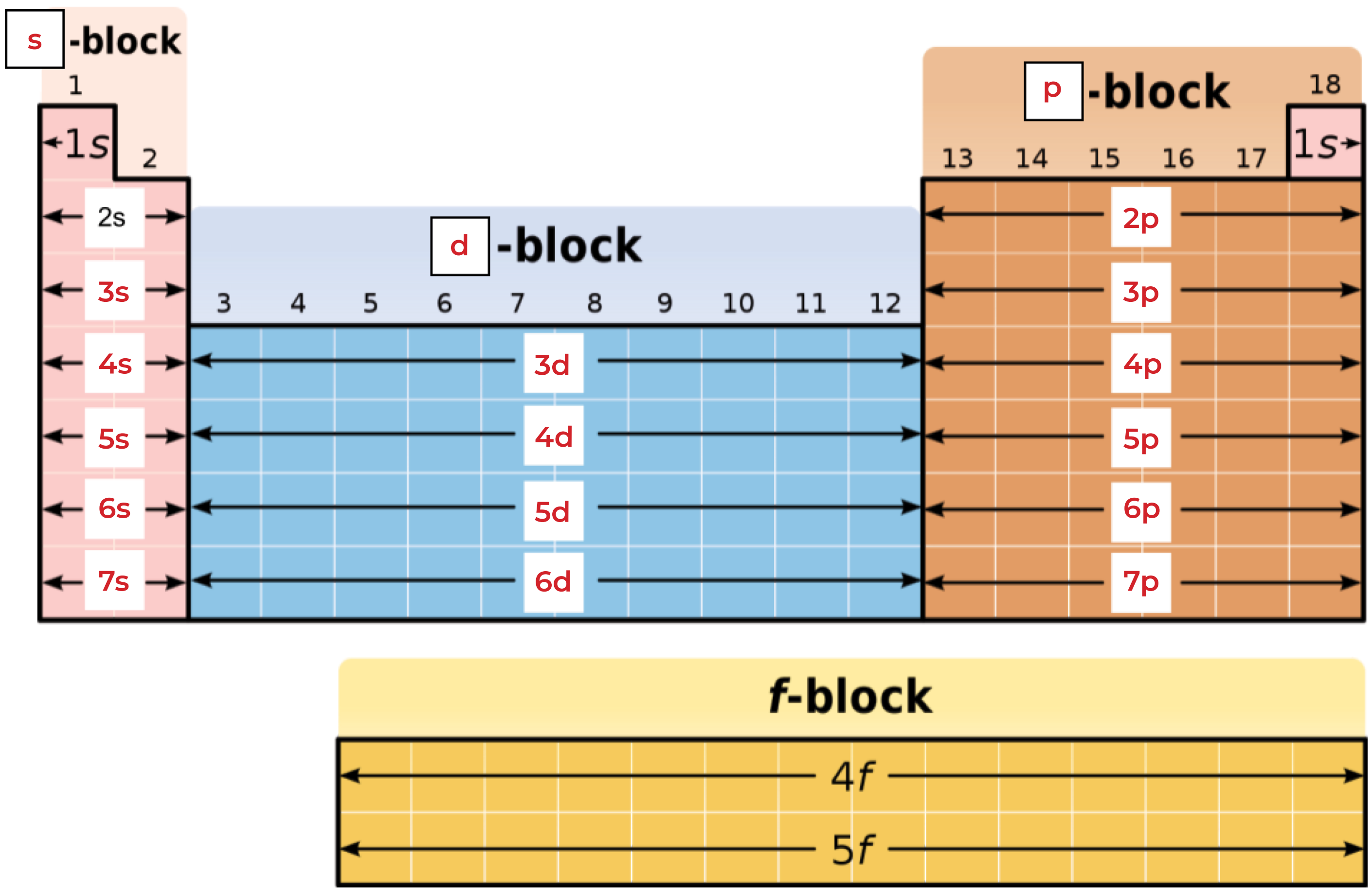
Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^3$

Noble Gas Configuration: $[Ne]\ 3s^2 3p^3$

OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

TASK 3: Using what you know about electron configurations, fill in the blanks with the orbital in which the last electron for indicated elements will be located. Notice how many columns are highlighted in each color to give you an idea of how many electrons can fit in that sublevel. You must also label each "block" of the periodic table with the correct orbital letter. The f-block has already been completed.

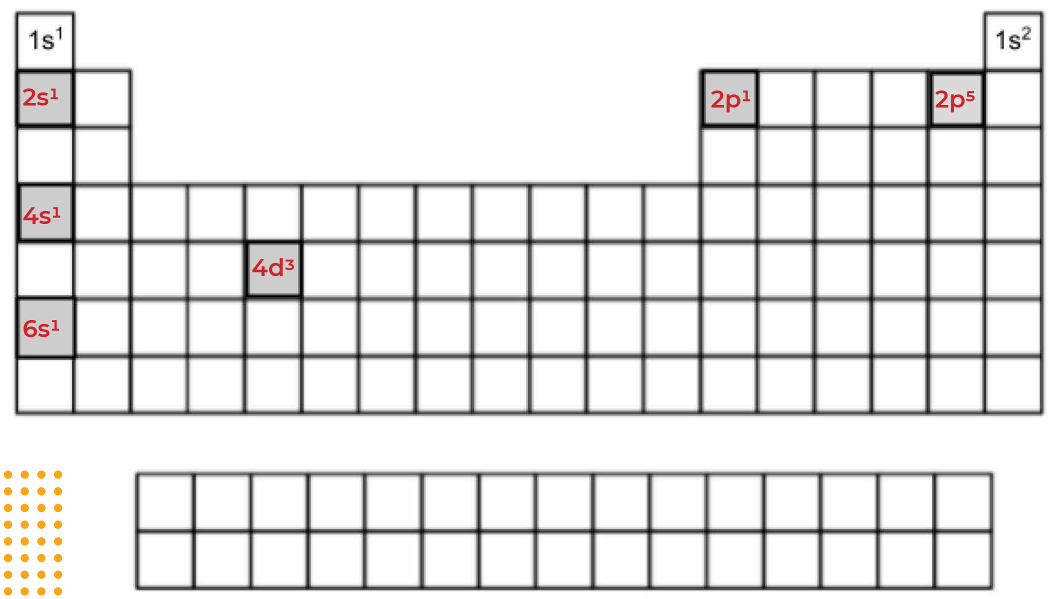


Periodic Table 2 by Roshan220195, CC BY-SA 3.0

OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

LOCK IT IN:
The periodic table can also be used to determine electron configuration. Fill in the highlighted spaces on the periodic table with the location of the last added electron in a ground state atom of that element. You will notice that you have already created some in the sandbox. However, there are five new ones you will also need to create. Hydrogen and helium have been filled in so that you can see the proper formatting. Pay careful attention to the energy level in which you are putting your final electron while working in the sandbox.



OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

TASK 4: Begin building an iron atom (Z=26) in the sandbox. Notice that electrons are numbered as they are added to the atom. As you add electrons 3, 5, 11, 13, 19, 21, and 26, identify the following for each:

- the portion of iron's electron configuration that represents that electron (e.g. the first electron would be $1s^1$)
- the principal quantum number (n)
- the angular momentum quantum number (l)
- all possible magnetic quantum numbers (m_l)

Do not forget to use your textbook and/or other resources to help guide you as needed!

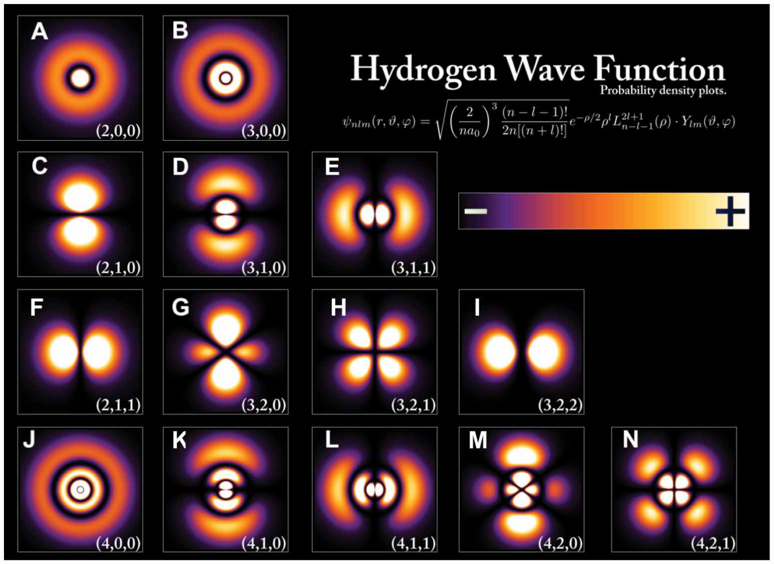
Electron Number in Sandbox	Corresponding Segment of Electron Configuration	Principal Quantum Number (n)	Angular Momentum Quantum Number (*l*)	Possible Magnetic Quantum Numbers (m_l)
1	$1s^1$	1	0	0
3	$2s^1$	2	0	0
5	$2p^1$	2	1	-1, 0, +1
11	$3s^1$	3	0	0
13	$3p^1$	3	1	-1, 0, +1
19	$4s^1$	4	0	0
21	$3d^1$	3	2	-2, -1, 0, 1, +2
26	$3d^6$	3	2	-2, -1, 0, 1, +2

OBJECTIVE 2

Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

TASK 5: Below are the probability density plots made from the wave functions for an electron in a hydrogen atom. Each is labeled with quantum numbers. Use this information to identify the orbital being represented by that function. Do not include orbital orientation. Once you identify the orbitals, compare the probability density plots in the diagram with the orbitals shown in the sandbox.

Image Letter	Quantum Numbers	Orbital
A	(2,0,0)	2s
B	(3,0,0)	3s
C	(2,1,0)	2p
D	(3,1,0)	3p
E	(3,1,1)	3p
F	(2,1,1)	2p
G	(3,2,0)	3d
H	(3,2,1)	3d
I	(3,2,2)	3d
J	(4,0,0)	4s
K	(4,1,0)	4p
L	(4,1,1)	4p
M	(4,2,0)	4d
N	(4,2,1)	4d



LOCK IT IN:
Based on your comparison of the probability density functions and the orbitals in the game, what does an orbital tell us?

An orbital as shown in the game is simply the region around the nucleus of an atom in which an electron with certain quantum numbers has the highest probability of being located.

OBJECTIVE 2

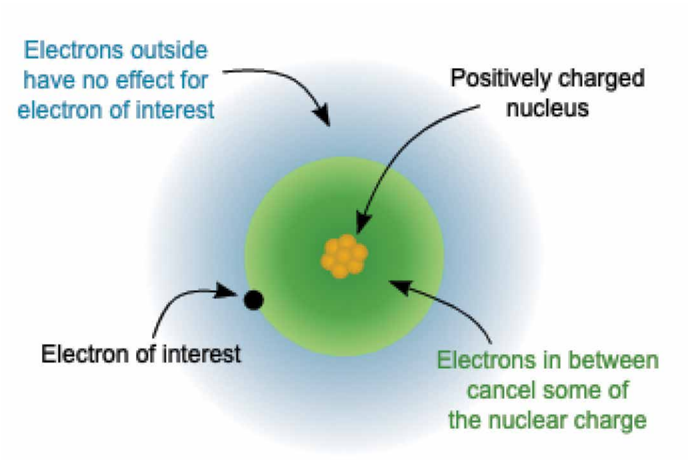
Demonstrate an understanding of how electrons are organized around the nucleus of an atom.

LOCK IT IN:
Construct the atoms in the table below in the sandbox using the information provided to you. Fill in any missing information.

Electron Configuration	Noble Gas Configuration	Quantum Numbers of Last Added Electron (n, l, m_l)	Element Name	Atomic Symbol	Atomic Number	Proton Number	Electron Number
$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$	(3,1,-1, -½) or (3,1,-0, -½) or (3,1,+1, -½)	Chlorine	Cl	17	17	17
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	$[Ar]4s^2$	(4,0,0, -½)	Calcium	Ca	20	20	20
$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$	3,0,0,+½	Sodium	Na	11	11	11

OBJECTIVE 3

Demonstrate an understanding of how effective nuclear charge determines trends in atomic radii.



TASK 6: Read the following information to help you determine the periodic trend.

Since all electrons share the same negative charge, repulsive forces help keep them generally separated. This effect is seen with "shielding", which occurs when core (non-valence) electrons reduce the **effective nuclear charge** or the attractive forces of the protons on the valence electrons. These electrons closer to the nucleus weaken the pull of the protons on the valence electrons and thus allow them to be more easily removed. Accordingly, increasing numbers of protons increases the effective nuclear charge on the valence electrons with only minor additional shielding as long as the added electrons are still in the same energy level. Addition of electrons in new energy levels increases shielding and thus reduces the attractive forces of the protons on the valence electrons even further.

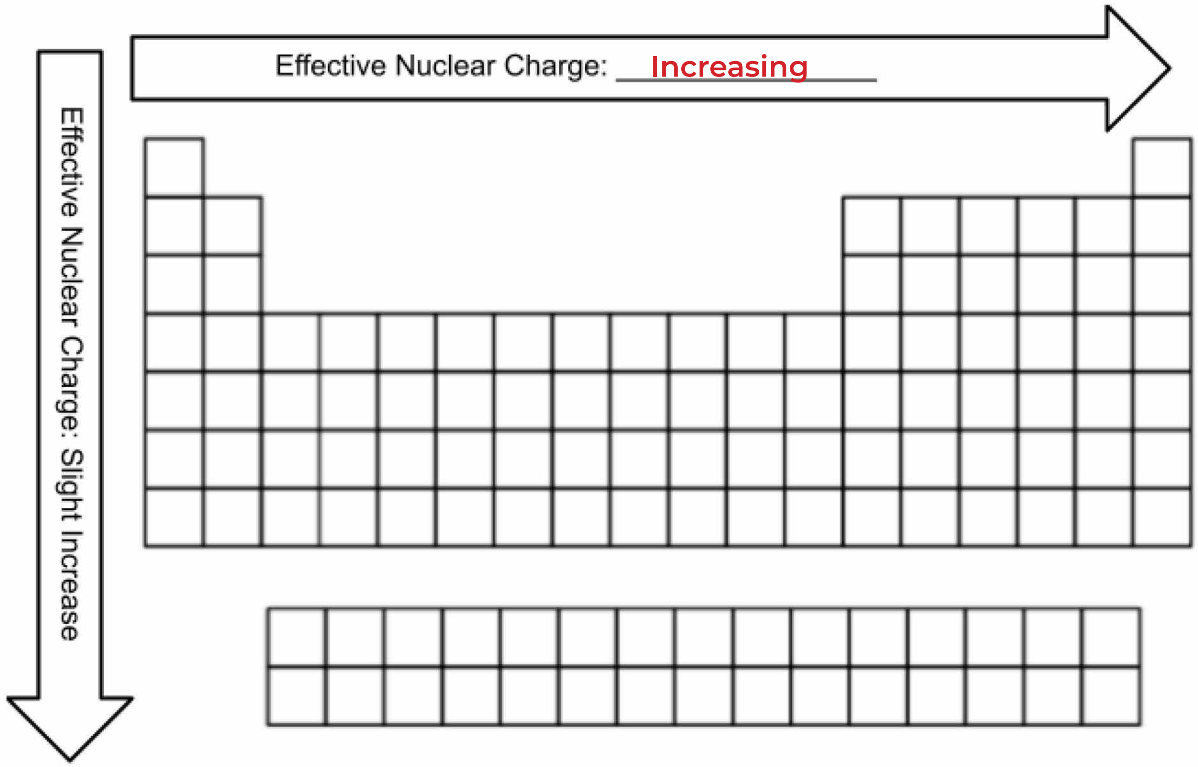
Consider the concept using the image of periods 4 and 5 of the periodic table below. Of the indicated elements, krypton (Kr) with its 35 protons has a larger effective nuclear charge on its valence electrons than chromium (Cr) which has a large effective nuclear charge on its valence electrons than potassium (K). The effective nuclear charge drops dramatically from krypton (Kr) to rubidium (Rb) despite the increased nuclear charge because the valence electrons are in a new energy level. However, note that in going down Group 1, the effective nuclear charge on the valence electrons of rubidium is slightly higher than that of potassium.

potassium 19 K 39.098	calcium 20 Ca 40.078	scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80
rubidium 37 Rb 85.468	strontium 38 Sr 87.62	yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29

OBJECTIVE 3

Demonstrate an understanding of how effective nuclear charge determines trends in atomic radii.

Use the information from the previous text to identify the trend in effective nuclear charge as you go right across a period by filling in the arrow below with the term "increasing" or "decreasing".



OBJECTIVE 3

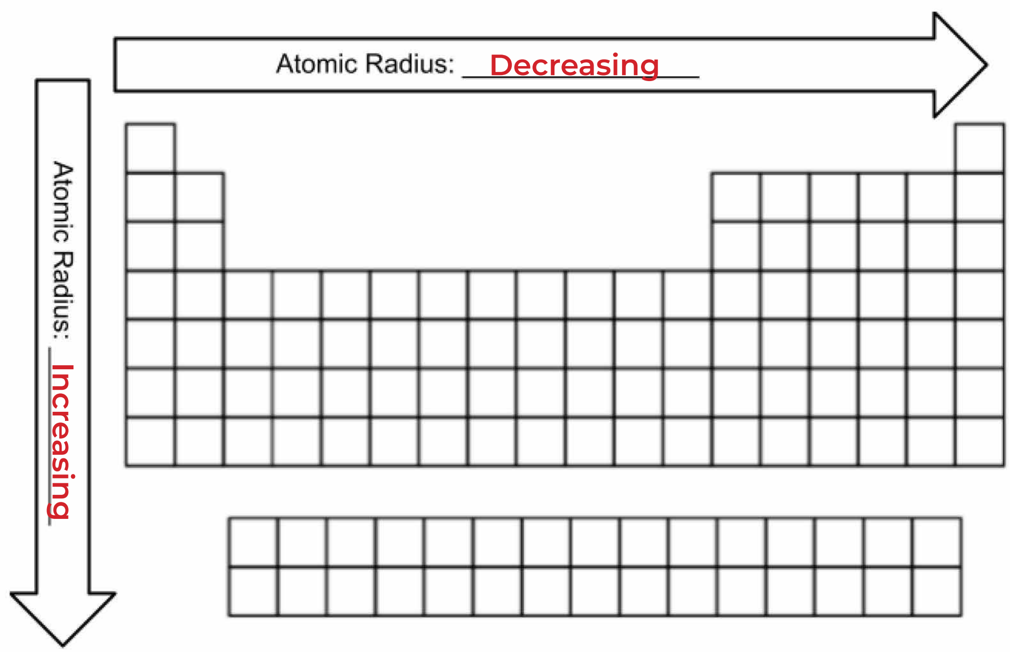
Demonstrate an understanding of how effective nuclear charge determines trends in atomic radii.

TASK 7: Take a look at the six original atoms that you made in Task 1. Place these six elements in order of increasing atomic radius in the spaces below.

LOCK IT IN:
Identify the trend in atomic radii as you go right across a period and down a group by filling in the arrows below with the terms "increasing" or "decreasing".

Fluorine < Boron < Copper < Lithium < Potassium < Cesium

Smallest Radius … Largest Radius



LOCK IT IN:
Use the terms **"energy level"** and **"effective nuclear charge"** to explain why the trend you saw in atomic radius going right across a period is different from the one in going down a group.

As one goes from left to right across a period, the number of protons (and thus the positive charge) increases. Although the number of electrons also increases with the addition of these protons, the electrons are all added in the same energy level. As a result, these electrons cause minimal shielding of the attractive force of the protons on the valence electrons. This results in an increase in effective nuclear charge going left to right across a period and a resulting decrease in atomic radius as the valence electrons are pulled closer to the nucleus. However, the addition of electrons at entirely new energy levels farther from the nucleus as you go down a group simply results in a larger atomic radius.

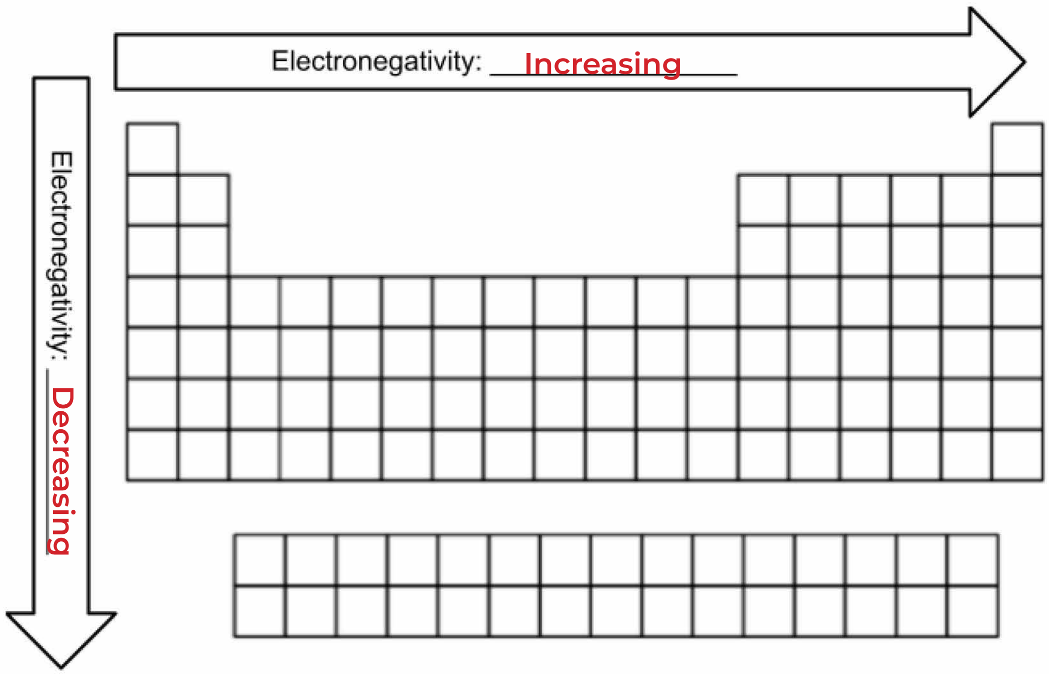
OBJECTIVE 4

Demonstrate an understanding of the periodic table trends in electronegativity.

TASK 8: Take a look at the six original atoms that you made in the sandbox. You should notice a red glow around them that varies in strength. This glow represents the **electronegativity** of the atom. Electronegativity describes an atom's tendency to attract electrons when it is bonded to another atom. It will become very important in other aspects of chemistry, but shows a clear periodic table trend like the ones you have seen in this sandbox activity. Order these six elements in order of increasing electronegativity in the spaces below.

Cesium < Potassium < Lithium < Copper < Boron < Fluorine

Least Electronegative ... Most Electronegative



LOCK IT IN:
Identify the trend in electronegativity as you go right across a period and down a group. Fill in the arrows with the terms "increasing" or "decreasing". Note that the last period does not follow this trend.

CLOSURE

CLOSURE: Compare the two atoms below (chlorine and magnesium) on their proton numbers, number of electrons, atomic radii, effective nuclear charge, and electronegativities using only a periodic table. Enter a greater than (>) or less than (<) symbol into the table. Then provide a brief justification as to why you chose the symbol you did using what you have learned. Once you have done so, create the atoms in the Sandbox to confirm your comparisons.

17	12
Cl	Mg

	> or <	Justification	Prediction Correct? (Y/N)
Proton Number	>	Chlorine has a higher atomic number.	YES
Number of Electrons	>	Chlorine has a greater number of protons.	YES
Atomic Radius	<	Magnesium is well to the left of chlorine in the same period.	YES
Effective Nuclear Charge on Valence Electrons	>	Chlorine is well to the right of magnesium in the same period.	YES
Electronegativity	>	Chlorine is well to the right of magnesium in the same period.	YES